

T05D05 – 5.3 Hess's Law of Summation

Name KEY

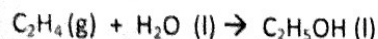
Hess' law is also called Hess' Law of Summation as you are adding reactions together in order to come up with the reaction you really want. This is done for several reasons. One is if you do not have a thermodynamics table of enthalpies, and another is if the reaction will not go to completion as it is written, and so I am unable to do the reaction in the lab.

Hess' Law states that it makes no difference whether I do the reaction in a single step or a series of steps the enthalpy must be the same value. [This is the 3rd rule of Thermochemistry]

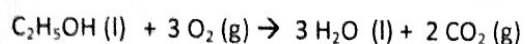
Let me illustrate with two examples.

1st example:

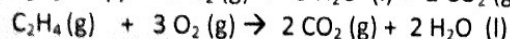
I want to calculate the enthalpy change for the following reaction:



and I am given the following reactions



$$\Delta H^\circ = -1367 \text{ kJ}$$

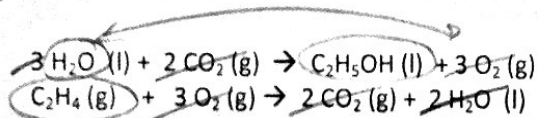


$$\Delta H^\circ = -1411 \text{ kJ}$$

The first thing I want to do is to determine what I must do. The first reaction I am given has $\text{C}_2\text{H}_5\text{OH}$ as a reactant, and I want it as a product therefore I will flip the first reaction and when I do that I must change the sign of the enthalpy $+1367 \text{ kJ}$ instead of -1367 kJ .

Next I notice that the second reaction has C_2H_4 as a reactant so I will keep that reaction the way it is.

Now I have



$$\Delta H^\circ = +1367 \text{ kJ}$$

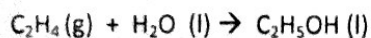
$$\Delta H^\circ = -1411 \text{ kJ}$$

Now I can add the two equations up (and hopefully they will be the equation at the top that I wanted to get).

When I add them up the two equations (just like adding simultaneous equations in math) the three oxygen's (3O_2) can be subtracted away (reactant on one side and product on the other side in equal numbers), and the two carbon dioxides (2CO_2) can also be subtracted. Two of the waters also can be eliminated ($3\text{H}_2\text{O}$'s on the reactant side and $2\text{H}_2\text{O}$'s on the product side).

So what I get is $\rightarrow \text{C}_2\text{H}_4(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{C}_2\text{H}_5\text{OH}(\text{l})$ ✓

This is the equation I wanted so now I add the two enthalpy values, and I get -44 kJ ($+1367 \text{ kJ}$ and a -1411 kJ), and I add this component to the above and make it a thermo chemical equation.



$$\Delta H^\circ = -44 \text{ kJ}$$

If I check this with thermo tables, you will get an answer very close to if not exactly the one listed above.

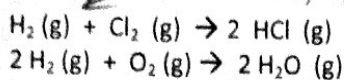
That is the beauty of Hess' law which states that if a reaction can be described by a series of steps, the ΔH (change in enthalpy) for the reaction is simply the sum of the enthalpies for each of the steps.

If I did the above equation in three steps or more instead of two the ΔH for the reaction would still be the same answer.

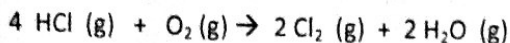
** Does not matter how many steps you take so long as the materials all cancel out to produce the same equation!*

Let's try another one:

Given the following data:

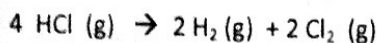


$$\begin{aligned} \Delta H^\circ &= -185 \text{ kJ} \quad \times 2 \\ \Delta H^\circ &= -483.7 \text{ kJ} \end{aligned}$$

I want to calculate ΔH° for the reaction

The first thing I should notice is that the first equation above has to be flipped and also doubled because HCl in the reaction I WANT has a coefficient of 4 in front of it. The second equation will be left alone. Do you see why? Oxygen is needed as a reactant and water is needed as a product. I am not going to double it as I only want 1 oxygen, and 2 waters.

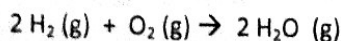
So this is what I get:



$$\Delta H^\circ = +370 \text{ kJ}$$

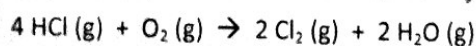
I have doubled the equation, and flipped it so therefore the ΔH value must also be doubled, and the sign must be change (I flipped the equation)

Now the second equation as written:



$$\Delta H^\circ = -483.7 \text{ kJ}$$

Now I add the two equations and get:

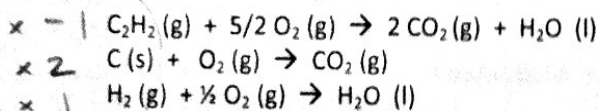


This is the equation I wanted, and now all I do is add the two enthalpies (ΔH° 's) and I get -113.7 kJ ($+370 \text{ kJ} - 483.7 \text{ kJ}$)

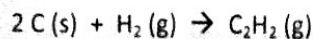
These take practice and patience. You have to plan before you act. You must determine which if any of the equations have to be flipped, and which one or ones have to be multiplied (or even divided). Once you get the equations set up correctly all you do is add up the enthalpies (ΔH° 's).

Now you try one:

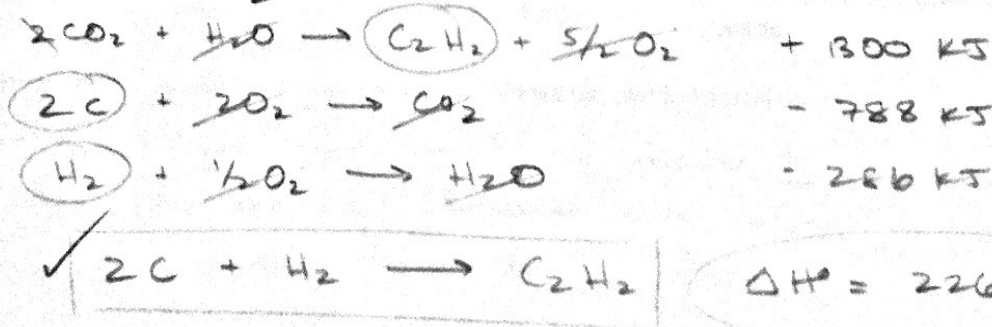
Given the following data:



$$\begin{aligned} \Delta H^\circ &= -1300 \text{ kJ} \\ \Delta H^\circ &= -394 \text{ kJ} \\ \Delta H^\circ &= -286 \text{ kJ} \end{aligned}$$

Calculate ΔH° for the reaction:

Three equations are more difficult than two but try it in the space provided.



When finished you should get an answer close to +227 kJ. If you got that answer good job. If not, try harder. Persistence pays off. These take practice. The only way you become good at anything is with practice.

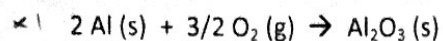
Remember you may have to double or triple an equation and also may have to flip the same equation. Take your time.

Some are easy and some are very difficult. They all take patience. Here are some more to try:

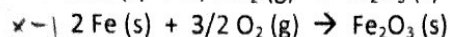
1. Calculate the standard enthalpy change for the reaction



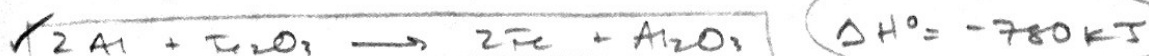
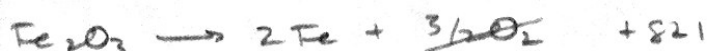
given that:



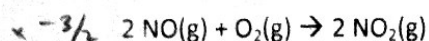
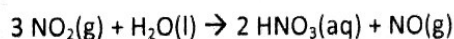
$$\Delta H^\circ = -1601 \text{ kJ}$$



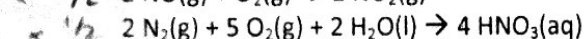
$$\Delta H^\circ = -821 \text{ kJ}$$



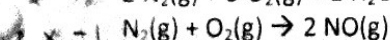
2. Calculate the standard enthalpy change (kJ) for the following reaction of nitrogen dioxide with water:



$$\Delta H^\circ = -114 \text{ kJ}$$



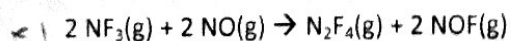
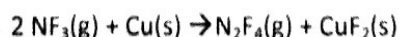
$$\Delta H^\circ = -255 \text{ kJ}$$



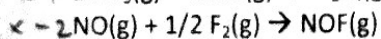
$$\Delta H^\circ = +181 \text{ kJ}$$



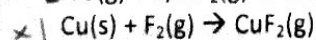
3. Given the following equations and ΔH° values, determine the heat of reaction (kJ) at 298 K for the reaction:



$$\Delta H^\circ = -82.9 \text{ kJ}$$



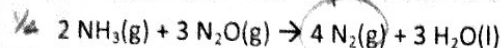
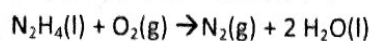
$$\Delta H^\circ = -156.9 \text{ kJ}$$



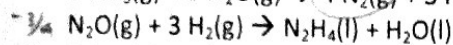
$$\Delta H^\circ = -531.0 \text{ kJ}$$



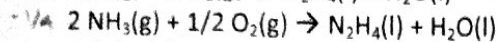
4. Given the following equations and ΔH° values, determine the heat of reaction (kJ) at 298 K for the reaction:



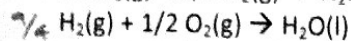
$$\Delta H^\circ = -1013 \text{ kJ}$$



$$\Delta H^\circ = -317 \text{ kJ}$$



$$\Delta H^\circ = -142.9 \text{ kJ}$$



$$\Delta H^\circ = -285.8 \text{ kJ}$$



$$-253.25 \text{ kJ}$$



$$+237.75 \text{ kJ}$$



$$+35.725 \text{ kJ}$$

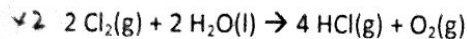
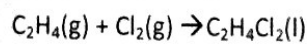


$$-643.05 \text{ kJ}$$



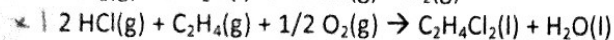
$$\Delta H^\circ = -622.825 \text{ kJ}$$

5. Determine ΔH° /kJ for the following reaction using the listed enthalpies of reaction:



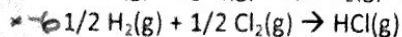
$$\Delta H^\circ = +202.5 \text{ kJ}$$

$$\times 1$$



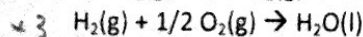
$$\Delta H^\circ = -319.6 \text{ kJ}$$

$$\times 1$$



$$\Delta H^\circ = -92.3 \text{ kJ}$$

$$\times -2$$

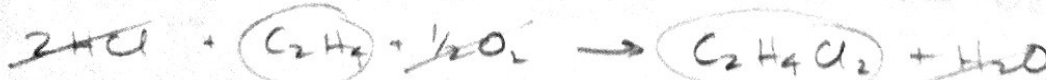


$$\Delta H^\circ = -285.8 \text{ kJ}$$

$$\times 1$$



$$+405.0 \text{ kJ}$$



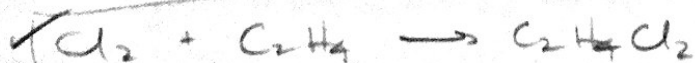
$$-319.6 \text{ kJ}$$



$$+553.8 \text{ kJ}$$



$$-857.4 \text{ kJ}$$



$$\Delta H^\circ = -218.3$$